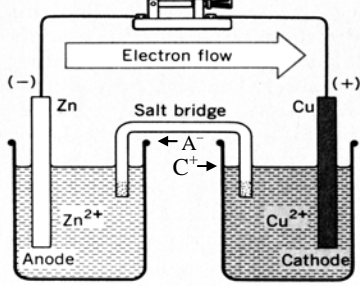


Galvanic/Voltaic	Electrolytic
1. Chemical \rightarrow Electric energy	1. Electric \rightarrow Chemical energy
2. Spontaneous redox reaction - produces electric current	2. Not spontaneous - requires electric supply
3. Two half cells are separated - connected through salt bridge	3. Both electrodes can be put in same container
4. Anode (-) Cathode (+) Oxidation = Anode Reduction = Cathode "An Ox, Red Cat"	4. Anode (+) Cathode (-) Oxidation = Anode Reduction = Cathode Still "An Ox, Red Cat" you just change the algebraic sign
5. e^- supplied by species being oxidized - move from anode to cathode	5. Power source (ex. Battery) supplies the e^- - go in through the cathode and come out through the anode
6. $\Delta G < 0, K > 1$	6. $\Delta G > 0, K < 1$

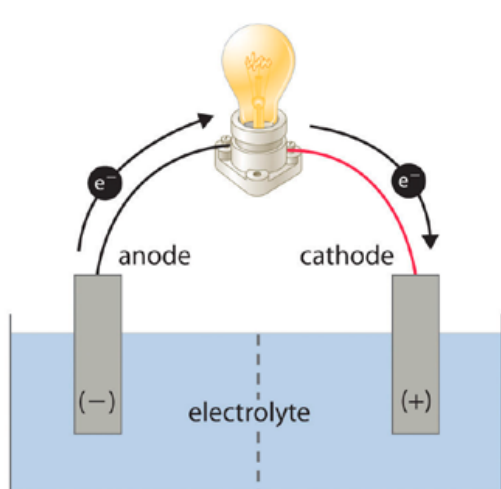
Note - Please check for typos! We pulled these pages together from quick things we found online. They seem good but haven't been proofread yet.

Explain why, in galvanic cells, the cathode is positive?	Galvanic cells spontaneously produce electricity. The negative electrode acquires electrons that are lost from metal atoms as they become positive ions. This is oxidation (LEO). Thus, the negative electrode is the anode (the anode, by definition is the site of oxidation). If the negative electrode is the anode, then the positive electrode must be the cathode.
Diagram the Zn-Cu galvanic cell. List all important aspects of the cell.	 <p>A^- = anion movement in salt bridge C^+ = cation movement in salt bridge</p> <p>Cell reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$</p>
What is a salt bridge? What is its function?	A salt bridge is a connection between the two half cells of a galvanic cell. It contains ions (e.g. K^+ and NO_3^-). The ions are essential to the function of the galvanic cell. As the half-cell ions decrease or build up, the ions of the salt bridge move into the half-cells to maintain electric neutrality.

Define cell potential.	The maximum emf (electrical force in volts) of a galvanic cell.
How can cell potential be increased?	By increasing the concentration of ions in half-cells, changing the composition of electrodes, or by connecting half-cells in series.
How can the direction of electron flow between half-cells be predicted?	By referring to a chart of standard reduction potentials.
What is meant by reduction potential?	The tendency of a half-cell to attract electrons (i.e. the tendency to be reduced).
What is meant by standard reduction potential? How is it symbolized?	The reduction potential of a half-cell under standard conditions (25°C, 1 atm, and ions at a concentration of 1 M). Symbol: E°
Why can't standard reduction potentials be measured? How is this problem overcome?	Half-cells do not conduct electricity until combined with another half-cell. Thus, standard reduction potentials are expressed relative to a reference electrode (the hydrogen/platinum electrode).
Give the half reaction for the hydrogen reference electrode.	$2H^+(aq) + 2e^- \leftrightarrow H_2(g)$
>What is the significance of \leftrightarrow	The reaction could proceed in either direction, depending on whether the reaction is oxidation or reduction. Note: it does not mean that an equilibrium exists.
What is a cell potential? Give its symbol.	A cell potential is the potential difference between two half-cells. In other words, it is the magnitude of difference between standard reduction potentials. Symbol: E°_{cell}
Give the equation for calculating cell potential.	$E^\circ_{cell} = (E^\circ \text{ of reduction}) - (E^\circ \text{ of oxidation})$
> What is an easy way to remember this equation?	$E^\circ_{cell} = (\text{larger } E^\circ) - (\text{smaller } E^\circ)$ In other words, E°_{cell} is always positive.

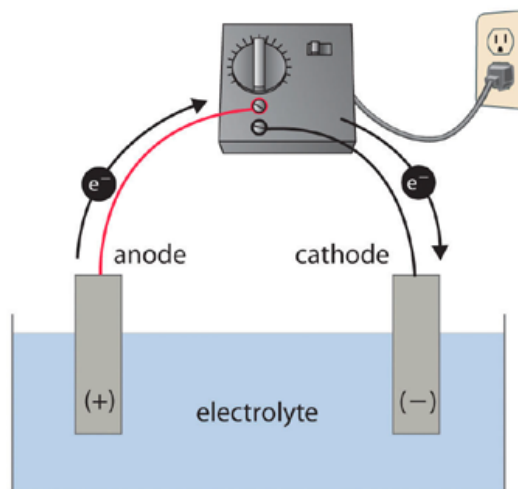
ELECTROCHEMICAL CELLS

Electrochemical Cells: A Comparison				
Galvanic (voltaic) cells	spontaneous oxidation-reduction reaction	Is separated into 2 half-cells	Electrodes made from metals (inert Pt or C if ion to ion or gas)	Battery – its cell potential drives the reaction and thus the e^-
Electrolytic cells	non-spontaneous oxidation-reduction reaction	Usually occurs in a single container	Usually inert electrodes	Battery charger – requires an external energy source to drive the reaction and e^-



GALVANIC CELL

Energy released by spontaneous redox reaction is converted to electrical energy.



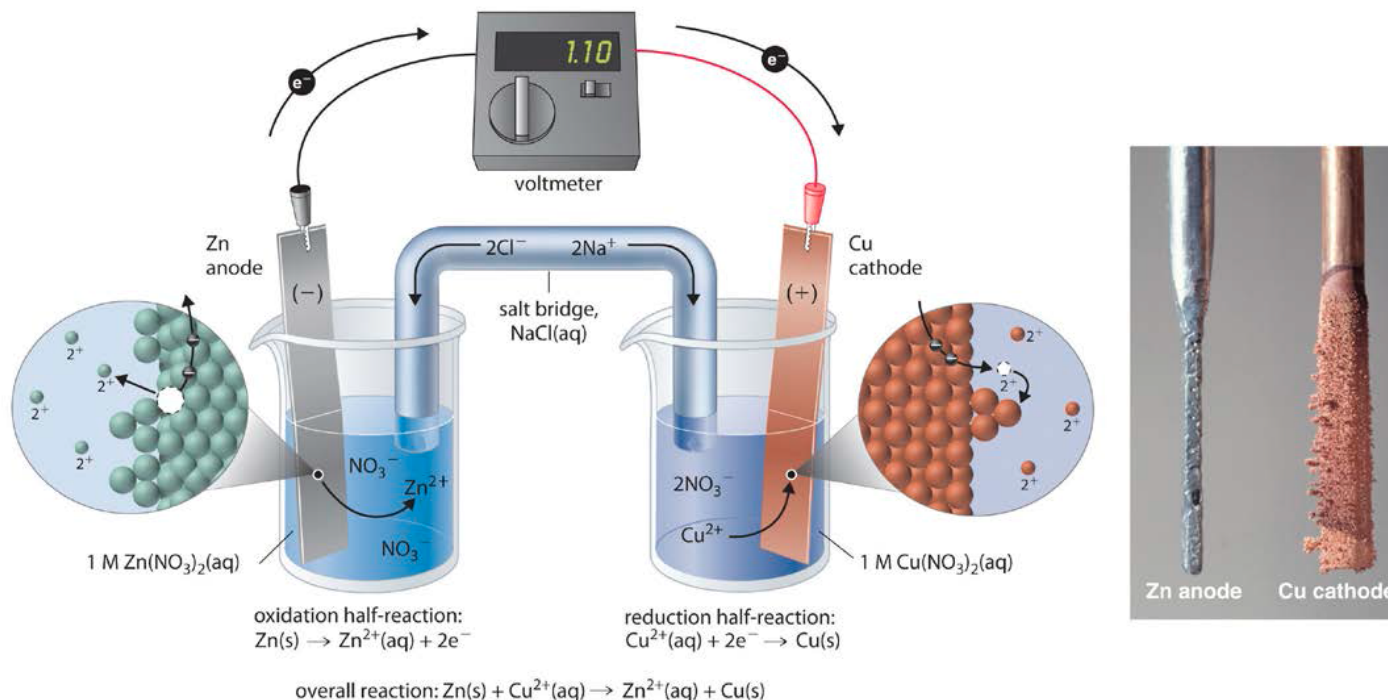
ELECTROLYTIC CELL

Electrical energy is used to drive nonspontaneous redox reaction.

The Galvanic Cell: *What is what and what to know?*

ANODE: AN OX	<u>Anode</u> – the electrode where oxidation occurs. Over time the mass of the anode may decrease as the metal is oxidized into ions.
CATHODE: RED CAT	<u>Cathode</u> – the electrode where reduction occurs. Over time the mass of the cathode may increase as the metal ions in the solution are reduced and plated onto it.
FAT CAT	Electron Flow – From Anode To CATHode
Ca + hode	Cathode is + galvanic cells
Salt Bridge	<u>Salt Bridge</u> – provides ions to balance the charge in each cell; contains a neutral salt that is very soluble (avoids precipitation issues). The salt cations flow into the cathode and the salt anions flow into the anode.

The Galvanic Cell: *How it Works!*



The above picture shows the oxidation-reduction reaction between Zn and Cu. The diagram of the cell clearly shows: the anode and cathode; the half-reaction that occurs at each electrode, the direction of electron flow, the direction of ion movement from the salt bridge, the overall reaction, as well as the E°_{cell} for the reaction.

Calculating Standard Cell Potential (E°_{cell})

The difference in electrical potential between the two half-reactions is measured with a voltmeter. The difference between the cell potentials of the two half-reactions determines the overall cell potential for the reaction.

1	Look at a table of Standard Reduction Potentials (or the reduction potentials that are provided). Write both reduction reactions from the table with their voltages.
2	THE MORE POSITIVE REDUCTION POTENTIAL IS REDUCED. THE LEAST POSITIVE IS OXIDIZED.
3	Reverse the equation that will be oxidized; be sure to change the sign of the voltage [this is now $E^{\circ}_{\text{oxidation}}$]
4	Balance and add the two half-reactions together.
5	Now add the two cell potentials together. $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{oxidation}} + E^{\circ}_{\text{reduction}}$
° indicates <i>standard conditions</i> (1 atm, 25°C; 1 M)	



More positive is reduced (+ 0.34 V > - 0.76 V) – Cu^{2+} (+0.34) is reduced and Zn oxidized

